Chemistry

Lecture 10 By: Syed Ali Shan

Fundamental Concepts

Outline:

- Atomic mass
- Empirical Formulae
- Molecular Formulae
- Mole
- Construction of mole ratios as conversion factors in stoichiometric calculation
- Avogadro's number
- Stoichiometry
- Important assumptions of stoichiometric calculations
- Limiting reactant
- Percentage yield

Relative Masses

Relative atomic mass (Ar):

- For elements
- Mass of an atom of an element compared to mass of atom of C-12
- H = 1.008 amu
- It is measured as average/fractional mass
- It's value is normally in fraction

Relative formula mass (Fr):

- For ionic compounds
- Sum of relative atomic masses atoms of an formula unit of an ionic compound
- NaCl = 58.5 amu

Relative molecular mass (Mr):

- For molecular/covalent compounds
- Sum of relative atomic masses atoms of a molecule of a covalent compound
- H_2O = 18 amu

Mass Number (A):

- For isotopes
- A = Z + N ⇒ Mass number = proton number + neutron number

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- Value is always a whole number
- ${}^{12}_{6}C \rightarrow$ mass number = 12

Isotopes (Not directly included in syllabus)

- Atoms of an element with same atomic numbers but different atomic masses (mass numbers)
- Phenomenon of isotopy given by Soddy who rejected Dalton's theory (atom is smallest particle of an element that can take part in reaction)
- Isotopes have same chemical properties (depend upon valence es^{-1})
- Isotopes have different physical properties (depend upon mass number)

Similarities of isotopes	Dissimilarities of isotopes
Same atomic number	Different mass number
Same proton number	Different number of neutrons
Same valence shell configuration	Different physical properties
Same chemical properties	Different half lives (radioactive)
Same place in periodic table	Different nuclear stability
Same symbol	Different relative abundance

Relative Abundance:

- It is the percentage of isotope in a mixture of isotopes of that element
- Properties of an element resemble with isotope of high relative abundance
- It is determined by mass spectrometry
- Total isotopes = 580
- Natural = 280 → Radioactive (unstable) = 40, Non-radioactive (stable) = 240
- Artificial = 300 (unstable, radioactive)
- Mono-isotopic ⇒ having 1 isotope i.e. **Arsenic (As), Gold (Au), Fluorine (F), Iodine (I)**
- C, H, O has 3 isotopes each, nickel has 5, calcium has 6, palladium has 6, cadmium has 9 and tin has 11 isotopes, N, Cl, Br 2 each and S has 4
- Silver has total 16 isotopes but 2 are stable
- Isotopes with even atomic number and mass number are more abundant
- Elements with odd atomic number almost never possess more than 2 stable isotopes
- 50% of earth's crust is made of ${}^{16}O$, ${}^{24}Mg$, ${}^{28}Si$, ${}^{40}Ca$ and ${}^{56}Fe$ (all multiple of 4)
- Out of 280 isotopes that occur in nature, 154 have even mass number and even atomic number and 86 with odd
- Physical methods for isotopes separation;
  Gaseous diffusion, thermal diffusion, distillation, centrifugation, electromagnetic separation, laser separation
- **Isobars** different atomic nuclei with same mass number i.e. ${}^{66}Zn$ and ${}^{66}Cu$
- **Isotones** different atomic nuclei with same number of neutrons i.e. ${}^{14}C$ and ${}^{16}O$
- **Isosteres** different molecules with same no. of atoms and valence electrons i.e. N_2O and CO_2

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Average/fractional atomic mass = $\frac{(\text{mass of 1st isotops} \times \text{it's abundance}) + (\text{mass of 2nd isotope} \times \text{it's abundance})}{100}$

It is the mass of an element that is obtained from isotopic mass and relative abundance of its isotopes.

Mole

- Fundamental SI unit for amounts of substances
- Atomic mass, molecular mass, formula mass and ionic mass of a substance expressed in grams
- [Atomic mass, molecular mass, formula mass and ionic mass] = Molar mass
- 1.008 g of hydrogen = 1 mole
- 18 g of water = 1 mole
- Mole = $\frac{\text{given mass of substance}}{\text{atomic mass/ molecular mass/ formula mass/ ionic mass}}$
- **n = m/M**
- **moles of atoms/ions/charges/electrons/etc in a substance = (m/M)× atomicity = n × atomicity**
- atomicity (atoms, ions, charges, electrons etc)

Example: 9 g of water (H_2O)

- $n_{H2O} = \frac{9}{18}$ = 0.5 moles
- moles of H-atom = m/M × atomicity = n × atomicity = 0.5 × 2 = 1 mole of H-atom
- moles of O-atom = n × atomicity = 0.5 × 1 = 0.5 mole of O-atom

Gram Atom:

- Atomic mass of an element expressed in grams.
- Gram atom = $\frac{\text{mass of an element}}{\text{atomic mass}}$
- 1.008 g of hydrogen = 1 gram atom

Gram Molecule:

- Molecular mass of a compound expressed in grams.
- Gram molecule = $\frac{\text{mass of compound}}{\text{molecular mass}}$
- 18 g of water (H_2O) = 1 gram molecule

Gram Formula:

- Formula unit mass of an ionic compound expressed in grams.
- Gram formula unit = $\frac{\text{mass of ionic compound}}{\text{formula unit mass}}$
- 58.5 g of NaCl = 1 gram formula

Gram Ion:

- Ionic mass of an ion expressed in grams.
- Gram ion = $\frac{\text{mass of an ion}}{\text{ionic mass}}$
- 17 g of OH^- = 1 gram ion

Avogadro's number

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- The number of atoms, molecules, formula units and ions present in one mole is called Avogadro's number.
- It is represented by N_A and its value is 6.02×10^{23}
- 1.008 g of (H) = 1 mole = 6.02×10^{23} atoms of H
- 18 g of (H_2O) = 1 mole = 6.02×10^{23} molecules H_2O
- Number of particles (atoms/molecules/formula units/ions) = $\frac{\text{mass of the substance}}{\text{atomic mass/ molecular mass/ formula mass/ ionic mass}} \times N_A$
- **No. of particles (N) = $\frac{m}{M} \times N_A$**
- **number of atoms/ions/charges/electrons/etc in a substance = $\frac{m}{M} \times N_A \times$ atomicity**

Or

- **number of atoms/ions/charges/electrons/etc in a substance = $n \times N_A \times$ atomicity**

Examples: 9 g of water (H_2O)

- No. of water molecules = $\frac{9}{18} \times 6.02 \times 10^{23}$
- No. of H-atoms = $\frac{m}{M} \times N_A \times$ atomicity = $\frac{9}{18} \times 6.02 \times 10^{23} \times 2$

Mixture of substances has 88 kg mass of 50% CO_2, molecules of CO_2?

- Mass = 88 × 1000 × 50/100 = 44000 g
- No. of CO_2 molecules = $\frac{44000}{44} \times 6.02 \times 10^{23}$

Molar Volume:

- Volume occupied by 1 mole of an ideal gas at STP
- It's value at STP is 22.414 dm^3 and at RTP is 24 dm^3
- 1 mole =Molar mass of substance(variable) = N_A (6.02×10^{23}) = 22.414 dm^3 (22414 cm^3) at STP
- **moles = $\frac{\text{Given/Reqiurd volume}}{\text{Volume at STP}}$ or $\frac{\text{Mass}}{\text{Molar mass}} = \frac{\text{Given/Reqiurd volume}}{\text{Volume at STP}}$**
- As we know;
- 1 mole of O_2 = 32 g of O_2 = 6.02×10^{23} molecules of O_2 = 22.414 dm^3 of O_2
  - What will be the volume of 16 g of O_2?
  - How many moles in 4 g of O_2?
  - How many N (molecules) in 64 g of O_2?
- Similarly it can be asked for any like how many moles in 14 g of N_2 etc

Empirical and Molecular Formulae

Empirical Formula (E.F)	Molecular Formula (M.F)
The simplest whole number ratio of atoms in a molecule	It gives actual number of atoms in a molecule
Percentage of element is required	E.F and molar mass are required
E.F = $\frac{M.F}{n}$	**M.F = n × E.F**
Used for ionic and covalent(molecular) compounds	For covalent(molecular) compounds
Benzene has CH, glucose has CH_2O	Benzene has C_6H_6, glucose has $C_6H_{12}O_6$
CH is empirical for acetylene and benzene, if n = 2 (acetylene) and n = 6 (benzene)	
CH_2O is empirical for acetic acid, formaldehyde, glucose and fructose, if n = 1 (formaldehyde) and n = 2 (acetic acid), n = 6 (glucose and fructose)	
Some compounds have same E.F and M.F i.e. H_2O, CO_2 etc	

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Determination of Empirical formula (only for numerical concept):

- Find the percentage composition of each element in the compound
- Find the number of gram-atoms (moles) of each element .For this purpose divide the percentage of each element by its atoms mass
- Find the atomic ratio of each element. To get this, divide the number of gram-atoms (Moles) of each element by the smallest number of gram-atoms (moles)
- Multiply with suitable to get whole number value if required

Combustion Analysis:

- Organic compound (having C,H,O) burnt in excess of oxygen
- Sole products are CO_2 and H_2O
- Determines empirical formula by providing %ages of elements
- Percentages of C, H are found directly
- Percentage of O by difference method (indirect)
- CuO is used to oxidize C completely to CO_2
- Water absorber is $Mg(ClO_4)_2$ → physical change
- CO_2 absorber is 50% KOH → chemical change

Stoichiometric Calculation

- Quantitative relationship between reactants and products
- Not for reversible reactions
- **Assumptions;**
  - All reactants converted to products
  - No side reactions
  - Law of conservation of mass and definite proportions being followed
- **Limitations of a chemical equation are;**
  - They do not tell about the conditions of reaction.
  - They do not give rate of reaction.
  - They can be written for a chemical change that actually does not occur

Stoichiometric Relationships:

1. **Mass-mass:**
   With the help of mass of given substance, mass of another substance can be calculated

- How many grams of CO_2 are produced by heating 50 g of $CaCO_3$?

$$CaCO_3 \rightarrow CaO + CO_2$$

$$100 \text{ g of } CaCO_3 \text{ gives } CO_2 = 44 \text{ g}$$

$$50 \text{ g of } CaCO_3 \text{ gives } CO_2 = 44/100 \times 50 = 22 \text{ g}$$

Or

$$\text{Moles of } CaCO_3 = 50/100 = 0.5 \text{ moles}$$

$$1 \text{ mole } CaCO_3 \text{ gives moles of } CO_2 = 1$$

$$0.5 \text{ mole } CaCO_3 \text{ gives moles of } CO_2 = 1/1 \times 0.5 = 0.5 \text{ moles of } CO_2$$

$$\text{Moles of } CO_2 = m/M$$

$0.5 = m/44 \qquad\qquad m = 22$ g

2. **Mass-mole:**

- With the help of mass of given substance, mole of another substance or vice versa can be calculated
- 1 mole of Ca is burnt in excess of O_2, how much CaO is produced?

$$2Ca + O_2 \rightarrow 2CaO$$

2 moles of Ca produces moles of CaO = 2

1 mole of Ca produces moles of CaO = $\frac{2}{2} \times 1 = 1$ mole

$$n = m/M$$

$$1 = \frac{m}{56} \qquad \rightarrow m = 56 \text{ g}$$

3. **Mole-mole:**

- With the help of mole of given substance, mole of another substance can be calculated
- 10 moles of N_2 produces moles of NH_3?

$$N_2 + 3H_2 \rightarrow 2NH_3$$

1 mole of N_2 produces moles of NH_3 = 2

10 mole of N_2 produces moles of NH_3 = $\frac{2}{1} \times 10 = 20$ moles

4. **Mass-volume:**

With the help of mass of given substance, volume of another substance or vice versa can be calculated

- How many dm^3 of CO_2 are produced by heating 50 g of $CaCO_3$?

$$CaCO_3 \rightarrow CaO + CO_2$$

Moles of $CaCO_3$ = 50/100 = 0.5 moles

1 mole $CaCO_3$ gives moles of CO_2 = 1

0.5 mole $CaCO_3$ gives moles of CO_2 = 1/1 × 0.5 = 0.5 moles of CO_2

Moles of CO_2 = volume required / Volume at STP

0.5 = volume / 22.414 dm^3

Volume = 11.2 dm^3

5. **Mole-volume:**

- With the help of mole of given substance, volume of another substance or vice versa can be calculated (will be solved as a step solved in mass-volume)

6. **Volume-volume:**

- With the help of volume of given substance, volume of another substance can be calculated
- How many cm^3 of CaO is produced when 11200 cm^3 of Ca is burnt in excess of oxygen?

$$2Ca + O_2 \rightarrow 2CaO$$

2 × 22414 cm^3 (2 moles from eq.) of Ca produces cm^3 of CaO = 2 × 22414 cm^3 (2 moles from eq.)

11200 cm^3 of Ca produces cm^3 of CaO = $\frac{2 \times 22414}{2 \times 22414} \times 11200 = 11200$ cm^3

Limiting Reactant:

- A reactant in smaller amount and control the amount of product formed
- Reactant in equation with higher coefficient is limiting reactant (**short cut**)

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- $2A + B \rightarrow C$ (A is limiting reacting)
- Burning of coal occurs in excess of oxygen. In this coal is limiting reactant
- Rusting of iron occurs in excess of oxygen present in air. So iron is limiting reactant
- Steps involved in determining limiting reactant;
  - ✓ Calculate the moles of each reactant
  - ✓ Calculate amount (mole) of product formed from each reactant using balanced chemical equation
  - ✓ The reactant that gives least amount (moles) of product is limiting reactant.

Yield:

- Amount of product being formed
- Actual (experimental) and theoretical (calculated)
- Actual always less than theoretical
- **%age yield** is calculated to find the efficiency of a chemical reaction.
- %age yield/ Efficiency = $\frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100$
- Actual yield is always less than parent atom
  - ✓ Reaction may be reversible
  - ✓ Some side reaction may occur
  - ✓ Due to mechanical loss (filtration, washing, drying etc)
  - ✓ Inexperience workers
  - ✓ Impurities may present
  - ✓ Miscalculation in measurement and calculation

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